CHAPTER 3

Atoms and molecules

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3.1 Introduction

The smallest unit of any element is called an atom. The atoms of different elements combine with one another to form new substances called compounds. The compounds so formed are neutral in character, i.e., they have no electric charges on them. All matter is made by the combination of atoms of different elements combined together in some fixed ratio. In a way, atoms are the basic building blocks of the matter.

There is a beautiful description of atomic theory in the **treaties**, "Vaisheshik Darshan" written by Kanad, in which he has proposed that the whole universe is made up of atoms.

By scientific study substance is classified into pure substance and mixture. Then substance is viewed as element and compound. Why one element differ from another? Similarly, why one compound is different from another? How new compounds are formed by the mutual action of elements or compounds?

By the late 1700s, scientists began to realize that the concept of atoms provided an explanation for many experimental observations.

3.2 Laws of chemical combination

One of the most important aspects of the subject of chemistry is the study of chemical reactions. These chemical reactions take place according to the certain laws, called the 'laws of chemical combination'.

(A) Law of conservation of mass

(B) Law of constant proportion

(C) Law of multiple proportion

(D) Law of reciprocal proportion

(E) Law of combining volumes (Gay Lussac's law of gaseous volume)

The laws of chemical combination are the experimental laws which led to the idea of atoms being the smallest unit of matter. The laws of chemical combination played a significant role in the development of Dalton's atomic theory of matter.

The first four laws deal with the mass relationships whereas the fifth law deals with the volumes of the reacting gases.

(1) Law of conservation of mass

This law is given by **Antoine Lavoisier**.

According to this law, matter is neither created nor destroyed in the course of chemical reaction although it may change from one form to other.

The total mass of materials present after a chemical reaction is the same as the total mass before reaction. The law was tested by Heydweiller and Landolt (1901-1906) who performed fifteen different reactions in a closed H-tube so that nothing was allowed to escape. It was noticed that there was hardly any change in the weights of the tube.

ACTIVITY 3.1

Consider any reaction, say combination of $AgNO_3$ and KI in H-tube (see figure). Let $AgNO_3$ and KI are filled in limb A and B, respectively and then tube is tilted to allow these reactants to react as $AgNO_3 + KI \rightarrow AgI + KNO_3$. The mass of tube before and after the reaction comes out to be same.

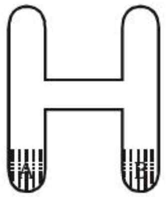

Fig. 1 H-tube

Example: 6.3 g of sodium bicarbonate was added to 15.0 g of acetic acid solution. Carbon dioxide gas was produced which was allowed to escape while the residue (of sodium acetate and water) was found to weigh 18.0 g. What is the mass of carbon dioxide which escaped out into the atmosphere?

Solution:

$NaHCO_3 + CH_3COOH \longrightarrow CH_3COONa + H_2O + CO_2$

According to law of conservation of mass,

Total mass before reaction = Total mass after reaction

$$\text{Mass of } NaHCO_3 + \text{Mass of } CH_3COOH = \underbrace{\text{Mass of } CH_3COONa + \text{Mass of } H_2O}_{18.0} + \text{Mass of } CO_2$$

$$6.3 \quad + \quad 15.0 \qquad\qquad 18.0 \quad + \text{ mass of } CO_2$$

Mass of CO_2 = 21.3 – 18.0 = 3.3 g

(2) Law of constant (or definite) proportion: [Proust, 1799]

According to this law, a chemical compound always contains the same elements combined together in the same proportion by mass.

This law was proved within the limits of experimental error by the work of Stas, who prepared the compounds in several different ways and showed that their composition was the same.

For example, CO_2 can be prepared by following ways:

(i) by heating $CaCO_3$: $CaCO_3 \xrightarrow{\Delta} CaO + CO_2$

(ii) by heating $NaHCO_3$: $2NaHCO_3 \xrightarrow{\Delta} Na_2CO_3 + H_2O + CO_2$

(iii) by burning carbon in O_2 : $C + O_2 \xrightarrow{\Delta} CO_2$

(iv) by $CaCO_3$ and HCl : $CaCO_3 + 2HCl \xrightarrow{\Delta} CaCl_2 + H_2O + CO_2$

CO_2 is collected separately as a product of each reaction and the analysis of CO_2 of each collection /reveals that it has the combination ratio of carbon and oxygen as 12: 32 or 3: 8.

Important aspects of law of constant proportion

(i) The discovery of isotopes, however, led to slight modification in this law. Since the isotopes of an element have different atomic weights, it is possible to have the same chemical compound with different composition according to the isotope used in its formation.

For example, CO_2 using C_{12} isotope has C: O:: 12: 32

CO_2 using C_{14} isotope has C: O:: 14: 32

(ii) Also the elements combining in the same ratio of their masses may give different compounds under different experimental conditions. For example, combination of carbon, hydrogen and oxygen in the ratio 12:3:8 may give C_2H_5OH or CH_3OCH_3 under different experimental conditions.

CONCEPTUAL CHALLENGE 3.1

Copper oxide was prepared by the following methods:

(a) In one case, 1.75 g of the metal were dissolved in nitric acid and igniting the residual copper nitrate yielded 2.19 g of copper oxide.

(b) In the second case, 1.14 g of metal dissolved in nitric acid were precipitated as copper hydroxide by adding caustic alkali solution. The precipitated copper hydroxide after washing, drying and heating yielded 1.43 g of copper oxide.

(c) In the third case, 1.45 g of copper when strongly heated in a current of air yielded 1.83 g of copper oxide. Show that the given data illustrate the law of constant composition.

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Explanation

In the first experiment,.

2.19 g of copper oxide contain 1.75 g of Cu.

∴ 100 g of copper oxide contains

$$= \frac{1.75}{2.19} \times 100 = 79.91\ \%$$

In the second experiment,

1.43 g of copper oxide contain 1.14 g of copper

∴ 100 g of copper oxide contains

$$= \frac{1.14}{1.43} \times 100 = 79.72\ \%$$

In the third experiment,

1.83 g of copper oxide contain 1.46 g of copper

∴ 100 g of copper oxide contains

$$= \frac{1.46}{1.83} \times 100 = 79.78\ \%$$

Thus, the percentage of copper in copper oxide derived from all the three experiments is nearly the same. Hence, the above data illustrate the law of constant composition.

(3) Law of multiple proportions (John Dalton 1804)

According to this law, when two elements A and B combine together to form more than one chemical compound then different weights of A, which combine with a fixed weight of B, are in a proportion of simple whole numbers.

The law is fully illustrated by the following example.

Nitrogen forms as many as five stable oxides.

That combines are in the ratio 16: 32: 48: 64: 90: i.e. 1: 2: 3: 4: 5 in N_2O, NO, N_2O_3, N_2O_4 and N_2O_5 respectively.

Example: Tin combines with oxygen to form two compounds having the following composition:

	% of Tin	% of oxygen
Compound A	78.77	21.23
Compound B	88.12	11:88

Show that the above data illustrates the law of multiple proportion.

Solution: In compound A:

21.23 parts of oxygen combine with 78.77 parts of tin

1 part of oxygen combines with 78.77/21.23 parts of tin = 3.7

In compound B:

11.88 parts of oxygen combine with 88.12 parts of tin

1 part of oxygen combines with 88.12/11.88 parts of tin = 7.4

Thus, weight ratio of tin for its combination with one part of oxygen are 3.7: 7.4 i.e. 1: 2. This is a simple ratio and this data illustrates the law of multiple proportion.

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CONCEPTUAL CHALLENGE 3.2

Carbon, combines with oxygen to form two compounds, namely, carbon dioxide and carbon monoxide. Show that law of multiple proportion is followed.

Explanation

In carbon dioxide, 12 parts by mass of carbon combine with 32 parts by mass of oxygen while in carbon monoxide, 12 parts by mass of carbon combine with 16 parts by mass of oxygen. Therefore, the masses of oxygen which combine with a fixed mass of carbon (12 parts) in carbon monoxide and carbon dioxide are 16 and 32 respectively. These masses of oxygen bear a simple ratio of 16: 32 or 1: 2 to each other.

(4) Law of equivalent proportion or law of reciprocal proportion: [Richter 1792-94]

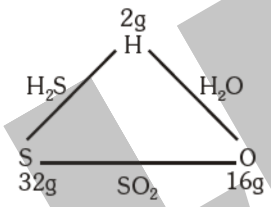
The weights of the two or more substances which separately react with same weight of a third substance are also the weights of these substances which react with each other or simple multiple of them. The law can be illustrated as below:

(a) Hydrogen combines with sulphur forming hydrogen sulphide (H_2S); 2 g of hydrogen reacts with 32 gm of sulphur.

(b) Hydrogen combines with oxygen forming water (H_2O); 2 g of hydrogen reacts with 16 g of oxygen.

(c) Sulphur combines with oxygen to form SO_2; 32 g of sulphur reacts with 32 g of oxygen i.e. in the ratio 32: 32. This ratio is double of the ratio of weights of these elements which combine with 2 g of hydrogen.

$$32/16 : 32/32 = 2 : 1$$



(5) Law of gaseous volumes: [Gay Lussac 1808]

According to this law, when gases combine, they do so in volume which bear a simple ratio to each other and also to the product formed provided all gases are measured under similar conditions.

The law can be illustrated as,

(a) $H_2(g) + \frac{1}{2} O_2(g) \longrightarrow H_2O(g)$

One volume of H_2 reacts with half volumes of O_2 to give one volume of H_2O vapours.

(b) $N_2 + 3H_2 \longrightarrow 2NH_3$

One volume of N_2 reacts with three volumes of H_2 to give two volumes of NH_3.

(c) $CO + \frac{1}{2} O_2 \longrightarrow CO_2$

One volume of CO combines with half volume of O_2 to give one volume of CO_2.

(d) $H_2 + Cl_2 \longrightarrow 2HCl$

One volume of H_2 combines with one volume of Cl_2 to give two volumes of HCl.

Example: 8 litre of H_2 and 6 litre of Cl_2 are allowed to react to maximum possible extent. Find out the final volume of reaction mixture. Suppose P and T remains constant throughout the course of reaction.

Solution:	H_2	+	Cl_2	$\rightarrow$	2HCl
Volume before reaction	8 litre		6 litre		
Volume after reaction	2		0		12 litre

Volume after reaction = Volume of H_2 left + Volume of HCl formed = 2 + 12 = 14 litre.

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3.3 Dalton's atomic theory

Keeping in view the laws of chemical combinations and the work of Greek philosophers, a meaningful atomic theory was finally proposed by an English school teacher John Dalton in 1803. The basic postulates of Dalton's theory are as follows:

(i) Each element is composed of extremely small particles called atoms.

(ii) All atoms of a given element are identical i.e. atoms of a particular element are all alike but differ from atoms of other elements.

(iii) Atoms of different elements posses different properties (including different masses).

(iv) Atoms are indestructible i.e. atoms are neither created nor destroyed in chemical reactions.

(v) Atoms of elements take part to form molecules i.e. compounds are formed when atoms of more than one element combine.

(vi) In a given compound, the relative number and kind of atoms are constant.

Advantages of Dalton's atomic theory

(i) Dalton's theory provides us a conceptual picture of matter. One can visualize an element as being comprised of tiny particles called atoms. Atoms are the basic building blocks of matter. They are the smallest units of an element that can combine with other elements in a chemical reaction. In compounds, the atoms of two or more elements combine in definite arrangements. Mixtures do not involve the intimate interaction between atoms that are found in compounds.

(ii) It is thus evident that Dalton's theory embodies several simple laws of chemical combinations that were known at that time. Postulate (iv) indicates for the law of conservation of mass. Postulate (vi) indicates the law of definite proportion.

(iii) Dalton's theory also explains the law of multiple proportion. It also explains that what makes an element to differ from each other.

Limitations of Dalton's atomic theory

(i) **Distinction between atoms and molecules:** According to Dalton, the smallest particle of an element as well as of compound was atom. However, he called 'atoms' the smallest particle of compound. Later on Avogadro used the term molecule for the 'compound atom' i.e. molecule is the smallest particle of compound.

(ii) **It could not justify Berzelius hypothesis:** Berzelius, a Swedish chemist, proposed that under similar conditions of temperature and pressure, equal volume of gases contain equal number of atoms. But Dalton's atomic theory could not explain this.

(iii) It could not explain why do atoms combine to form a molecule.

(iv) It could not explain the nature of forces which hold the atoms and molecules in solids, liquid and gaseous state.

(v) It could not explain the Gay Lussac's law of combining volume.

(vi) It could not explain that why atoms of an element differ in their masses.

3.4 Chemical symbols of elements

Before 1600 AD, alchemists tried to represent the substances that they used to their experiments by various kinds of pictorial symbols such as a triangle representing the earth for silver and so on.

Dalton used some other types of symbols to represent the elements. For instance, he drew a circle for oxygen or a circle with a dot for hydrogen.

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Later, Johann Berzelius, a Swiss chemist suggested that the initial capital letter should represent a particular element such as O for oxygen, H for hydrogen and C for carbon and so on. But in some cases, the symbols or the alphabet he had suggested did not agree with the English names of elements. This was because some of his symbols were based on the Latin names of the elements as well as certain other elements on their Greek names. But the method suggested by him forms the basis on which the modern system of chemical symbols and formulae are taken into account.

Symbols obtained from other languages:

Common name (of the element)	Symbol	Latin (L) or Greek (G) name
Gold	Au	Aurum (L)
Silver	Ag	Argentum (L)
Copper	Cu	Cuprum (L)
Lead	Pb	Plumbum (L)
Mercury	Hg	Hydrargyrum (L)
Iron	Fe	Ferrum (L)
Potassium	K	Kalium (L)
Tin	Sn	Stannum (L)
Nitrogen	N	Nitrogenium (L)
Carbon	C	Carbonium (L)
Oxygen	O	Oxygenium (L)
Silicon	Si	Silex (L)
Calcium	Ca	Clax (L)
Hydrogen	H	Hydrogenium (L)
Antimony	Sb	Stibium (L)
Sodium	Na	Natrium (L)
Chromium	Cr	Chromos (G)
Krypton	Kr	Kryptos (G)

A symbol is a short form that stands for the atom of an element. Each element is denoted by a symbol. This is usually the first letter of its name in English or Latin or Greek. For example, the symbol O represents the element oxygen, Latin name for copper is Cuprum, so the letters Cu represents copper. The symbol of an element also represents the mass of the element which contains one Avogadro's number of atoms of that element.

Thus,

H_2 represents one molecule of hydrogen.

H_2O represents one molecule of water.

Na_2CO_3 denotes one formula unit of sodium carbonate.

CH_3COOH denotes one molecule of acetic acid.

$CuSO_4.5H_2O$ denotes one formula unit of hydrated copper sulphate.

Importance of molecular formula

The molecular formula of a compound has quantitative significance. It represents:

(1) The representative numbers of different atoms in one molecule of the compound.

(2) The ratios of the respective masses of the elements present in the compound.

(3) The molecular formula represents one molecule of the substance, for example, CO_2 represents one molecule of carbon dioxide.

(4) Molecular formula gives the name and the actual number of atoms of all kinds present is one molecule of the substance. Thus, the formula CO_2, indicates that one molecule of carbon dioxide contains one atom of carbon and two atoms of oxygen.

Let us illustrate. The formulae CO_2

(a) This means that the molecular formula of carbon dioxide is CO_2.

(b) Each molecule contains one carbon atom joined by chemical bonds with two oxygen atoms.

(c) The molecular mass of carbon dioxide is 44 given that the atomic mass of carbon is 12 and that of oxygen is 16. (12+ 16+16 = 44)

Similarly,

1. H_2 is the molecular formula of hydrogen. This formula shows that one molecule of hydrogen contains two atoms of hydrogen.
2. H_2O is the molecular formula of water. This formula shows that one molecule of water contains two atoms of hydrogen and one atom of oxygen.

3.5 How to write the formula of compound

The molecular formula of a compound is electrically neutral, that is, the positive and the negative valencies of the ions or the radicals present in the compounds add up to zero.

The following steps should be taken while attempting to write a formula:

(a) Write the symbols side by side, usually the one with the positive valency first. The ion containing more than one atom should be written within brackets.

(b) Write the valency of each ion on the top of its symbol.

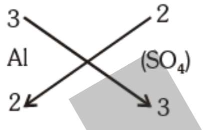
(c) Divide the valency numbers by their Highest Common Factor to get the simple ratio. Ignore the (+) or (–) signs of the radicals. If the valency number is one, it need not be written.

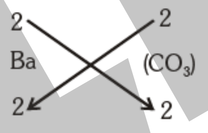
(d) Write the interchanged valency numbers to the lower right of the radicals, if the radical is a group of atoms and receives a valency number of more than one then enclose it within brackets.

Take the case of aluminium sulphate and barium carbonate.

Step 1: Al^{3+} $(SO_4)^{2-}$ Ba^{2+} $(CO_3)^{2-}$

Step 2:





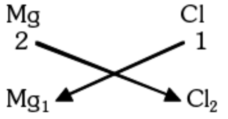
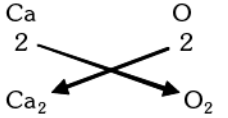
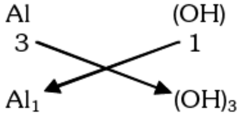
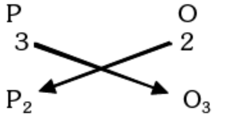
Step 3: Dividing the numbers by their HCF

$Al_2(SO_4)_3$, $Ba_1(CO_3)_1$

Step 4: Valency number 1 is not written

$Al_2(SO_4)_3$ $BaCO_3$

Hence, we can see that in $Al_2(SO_4)_3$ molecule, $3 \times 2 = +6$ charges on two Al^{3+} ions are balanced by $-2 \times 3 = -6$ charges on three SO_4^{2-} ions. Similarly, in $BaCO_3$, molecule, $2 \times 1 = +2$ charges on one Ba^{+2} ion balanced by -2 charges on one CO_3^{2-} ion.

Name of compound	**Symbols with valencies**	**Exchanges of valency**	**Formula**
Magnesium Chloride	Mg^{2+} Cl^{1-}		$MgCl_2$
Calcium Oxide	Ca^{2+} O^{2-}		CaO [cancelling common factor]
Aluminium Hydroxide	Al^{3+} $(OH)^{1-}$		$Al(OH)_3$
Phosphorus trioxide	P^{3+} O^{2-}		P_2O_3

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Naming of salts

Acids from which the salt is produced	Suffix and name of the salt
1. 'ous' acid	'ite'
Example:	
Sulphurous acid (H_2SO_3)	$CaSO_3$ Calcium sulphite
Nitrous acid (HNO_2)	$Zn(NO_2)_2$ Zinc nitrite
Phosphorous acid (H_3PO_3)	$Mg_3(PO_3)_2$ Magnesium Phosphite
2. 'ic' acid	'ate'
Example:	
Sulphuric acid (H_2SO_4)	$ZnSO_4$ Zinc sulphate
Nitric acid (HNO_3)	$NaNO_3$ Sodium nitrate
Phosphoric acid (H_3PO_4)	$AlPO_4$ Aluminium phosphate

The positive radical present in the salt comes from the corresponding base and the negative radical comes from the corresponding acid.

The name of the salts starts with the name of the metal present as positive radical which is followed by the name of the negative radical. The name of the negative radical is determined by the name of the acid from which the salt is produced.

Anions

(1) Monoatomic anions are named by adding 'ide' to the stem name of the element.

Examples are H^- hydride; F^- fluoride; Se^{2-} selenide; N^{3-} nitride; P^{3-} phosphide.

(2) Certain polyatomic anions have their names ending in 'ide':

OH^-	Hydroxide	CN^-	Cyanide	N_3^-	Azide
O_2^{2-}	Peroxide	S_2^{2-}	Disulphide	I_3^-	Triiodide
O_2^-	Superoxide	C_2^{2-}	Acetylide	NH^{2-}	Imide
O_3^-	Ozonide			NH_2^-	Amide

Ions such as HS^- and HF_2^- are named hydrogensulphide and hydrogendifluoride.

(3) **Oxyacid anions:** Oxyanions are named using the name characteristic of the central atom ending in 'ate', with the oxidation number of the central atom being shown in roman numerals put in parentheses after the name of the anion. Examples,

SO_4^{2-} Tetraoxosulphate (VI) SO_3^{2-} Trioxosulphate(IV)

PO_3^{3-} Trioxophosphate (III) $S_2O_3^{2-}$ Trioxothiosulphate (VI)

(4) Certain anions have well established trivial names in which prefixed such as per, hypo, etc. are used. This is in conformity with the name of the corresponding acids: the 'ous' acids give 'ite' anions and the 'ic' acids give the 'ate' anions. Examples are:

NO_3^-	Nitrate	$N_2O_2^{2-}$	Hyponitrite
SO_3^{2-}	Sulphite	ClO_2^-	Chlorite
SeO_3^{2-}	Selenite	$S_2O_5^{2-}$	Bisulphite
PO_3^{3-}	Phosphite	AsO_3^{3-}	Arsenite
NOO_2^-	Peroxonitrite	ClO_4	Perchlorate
$S_2O_2^{2-}$	Thiosulphite		

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(5) Polyatomic cations derived from the addition of protons to mono-atomic anions are named by adding-onium to the root name of the anion, e.g. phosphonium, arsonium, iodonium ions. Exceptions are ammonium, hydroxyl, hydrazinium, anilinium, pyridinium, etc.

The following trivial names are acceptable:

H_2O	Water	NH_3	Ammonia	BH_3	Borane
AsH_3	Arsine	SiH_4	Silane	SbH_3	Stibine
PH_3	Phosphine	Si_2H_6	Disilane	N_2H_4	Hydrazine

3.6 Empirical and molecular formula

The chemical formula of the compound that can be calculated from the element analysis data (percentage by mass of each element in the compound) is called the empirical formula.

Empirical formula represents the simplest whole number ratio of the atoms of different constituent elements present in one molecule of the compound.

Steps taken to arrive at the emperical formula

(i) Determine the percentage of each element: Percentage of each element is determined quantitatively. However, the percentage of oxygen is obtained by subtracting the sum of all other percentages from 100 i.e. % of O = [100 – sum of % of all other constituent elements].

(ii) Divide the percentage of each element by its atomic mass to obtain the relative numbers of atoms (or atomic ratio).

$$\text{Atomic ratio} = \frac{\text{Percentage of an element}}{\text{Atomic mass of the same element}}$$

(iii) Divide the atomic ratio by smallest quotient to get the simplest ratio of various elements.

(iv) Convert the simplest ratio to the whole number ratio either by,

  (1) Rounding them off to whole number, if the value is quite close to whole number. For example, 1.98, 2.99, 3.95 are rounded off as 2, 3 and 4 respectively.

  (2) Multiply the figures by suitable integer (2, 3 or 4, etc.)

(v) Write the empirical formula of the compound by writing the symbols of the various elements side by side. Now insert the whole number ratio of each element as the subscripts to the lower right hand corner of each symbol.

CONCEPTUAL CHALLENGE 3.3

An inorganic salt gave the following percentage composition Na = 29.11, S =40.51 and O = 30.38. Calculate the empirical formula of the salt.

Explanation

Calculation of empirical formula

Element	Symbol	Percentage of element	At. mass of elements	Moles of the element = Percentage At. mass (Relative no. of moles)	Simplest molar ratio	Simplest whole no. molar ratio
Sodium	Na	29.11	23	$\frac{29.11}{23} = 1.266$	$\frac{1.266}{1.266} = 1$	2
Sulphur	S	40.51	32	$\frac{40.51}{32} = 1.266$	$\frac{1.266}{1.266} = 1$	2
Oxygen	O	30.38	16	$\frac{30.38}{16} = 1.897$	$\frac{1.89}{1.266} = 1.5$	3

Thus, the Empirical formula is $Na_2S_2O_3$.

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Molecular formula

It is that formula of the compound which gives the actual number of atoms of various elements in a molecule of that compound. For example, molecular formula of glucose is $C_6H_{12}O_6$. Molecular formula of a compound is related to empirical formula as being a simple whole number multiple of empirical formula.

The molecular and empirical formula of any chemical compound are related as n ×Empirical formula = Molecular formula, where 'n' is positive integer.

$$n = \frac{\text{Molecular mass of compound}}{\text{Empirical formula mass}}$$

For example, Let us consider hydrogen peroxide whose empirical formula is HO and its molecular formula mass is 34.

$$N = \frac{\text{Molecular formula mass of hydrogen peroxide}}{\text{Empirical formula mass of hydrogen peroxide}} = \frac{34}{1+16} = \frac{34}{17} = 2$$

Thus, the molecular formula of hydrogen peroxide = (empirical formula)$_n$ = $(HO)_2$ = H_2O_2.

When n = 1. Molecular formula = Empirical formula

The molecular mass of a volatile compound can be determined by Victor Meyer's method (based on the principle that 22.4L of vapours of a substance at STP have mass equal to gram molecular mass) or by employing the relation.

Molecular mass = 2 ×Vapour density.

Empirical formula mass can however, be obtained from its empirical formula simply by adding the atomic masses of the various atoms present in it. Thus, for glucose (E.F. = CH_2O) the empirical formula mass can be calculated as follows:

Empirical formula mass = At. mass of carbon + 2 × At. mass of hydrogen + At. mass of oxygen

= 12.0 u + 2 × 1.0 u + 16.0 u = 30.0 u.

Hence, following steps have to be carried out to arrive at the molecular formula of the compound.

1. Determine the empirical formula mass by adding the atomic masses of all the atoms represented in the empirical formula. For example, empirical formula of glucose is CH_2O.
   ∴ Empirical formula mass = 12 + 2 × 1 + 16 = 30u
2. Determine the molecular mass from the data obtained by a suitable experiment.
3. Divide the molecular mass by the empirical formula mass and find out the value of n.
4. Multiply empirical formula by n to get the molecular formula.

CONCEPTUAL CHALLENGE 3.4

An organic substance containing carbon, hydrogen and oxygen gave the following percentage composition.

C = 40.68%; H = 5.085% and O = 54.228%

The vapour density of the compound is 59. Calculate the molecular formula of the compound.

Explanation

Calculate the empirical formula of the compound.

Element	Symbol	Percentage of element	At. mass of elements	Moles of the element = Percentage At. mass (Relative no. of moles)	Simplest molar ratio	Simplest whole no. molar ratio
Carbon	C	40.687	12	$\frac{40.687}{12} = 3.39$	$\frac{3.390}{3.389} = 1$	2
Hydrogen	H	5.085	1	$\frac{5.085}{1} = 5.085$	$\frac{5.085}{3.389} = 1.5$	3
Oxygen	O	54.228	16	$\frac{54.228}{16} = 3.389$	$\frac{3.389}{3.389} = 1$	2

Empirical formula is $C_2H_3O_2$.

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Calculate the empirical formula mass.

Empirical formula mass = $(2 \times 12) + (3 \times 1) + (2 \times 16) = 59$.

To calculate the molecular mass of the salt. The vapour density of the compound = 59. (Given)

Using the relation between vapour density and molecular mass,

Molecular mass = 2 × vapour density = $2 \times 59 = 118$.

To calculate the value of 'n'

$$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}} = 2$$

To calculate the molecular formula of the salt we use the following formula.

Molecular formula = n ×Empirical formula = $2 \times C_2H_3O_2 = C_4H_6O_4$

Thus, the molecular formula is $C_4H_6O_4$.

3.7 Mole concept and Avogadro's hypothesis

There are many ways of measuring the amount of a substance, weight and volume being the most common. But the basic unit of chemistry is the atom or a molecule and to measure the number of atoms or molecules is, therefore, of foremost importance.

Mole in Latin means heap or mass or pile. A mole of atoms is a collection of atoms whose total weight is the number of grams equal to the atomic weight. As equal number of moles of different elements contain equal number of atoms, it is convenient to express amounts of the elements in terms of moles. Just as a dozen means twelve objects, a score means twenty objects, chemists have defined a mole as a 'definite number' of particles, viz., atoms, molecules, ions or electrons, etc. This 'definite number' is called the **Avogadro constant**, equal to 6.023×10^{23}, in honour of Amedeo Avogadro.

The value of the Avogadro constant depends on the atomic-mass scale. At present the mole is defined as the amount of a substance containing as many atoms, molecules, ions, electrons or other elementary entities as there are carbon atoms in exactly 12 g of ^{12}C.

The following are the definitions of 'mole' represented in the form of equations:

(i) Number of moles of molecules = $\frac{\text{Weight in g}}{\text{Molecular mass}}$

(ii) Number of moles of atoms = $\frac{\text{Weight in g}}{\text{Atomic mass}}$

(iii) Number of moles of gases = $\frac{\text{Volume at NTP}}{\text{Standard molar volume}}$

(Standard molar volume is the volume occupied by 1 mole of any gas at NTP, which is equal to 22.4 litres.)

(iv) Number of moles of atoms / molecules / ions / electrons = $\frac{\text{Number of atoms / molecules / ions / electrons}}{\text{Avogadro constant}}$

(iv) Number of moles of solute = Molarity ×Volume of solution in litres.

At same temperature and pressure, equal volumes of gases contain the same number of molecules. Thus, for homogeneous gaseous reactions, the stoichiometric coefficients of the chemical equation also signify the relative volumes of each reactant and product under the same conditions of temperature and pressure.

For example,	H_2 (g) +	I_2 (g)	$\longrightarrow$	2HI (g)
	1 molecule	1 molecule		2 molecules
	Or 1 mole	1 mole		2 moles
	Or 1 volume	1 volume		2 volumes (T and P constant)
	Or 1 pressure	1 pressure		2 pressures (T and V constant)

1 mole is a fixed number of particles but not a fixed weight.

If conservation of mass, is expressed in the concepts of atoms means the concept of moles then if atoms are conserved, moles of atoms shall also be conserved. This is known as the **principle of atom conservation**. This principle is in fact the basis of the mole concept.

Examples:

Q.1 Calculate the weight of 6.023×10^{23} molecules of $CaCO_3$.

Sol. No. of moles of $CaCO_3 = \frac{\text{Number of molecules}}{\text{Avogadro constant}}$

$$\frac{6.023 \times 10^{23}}{6.023 \times 10^{23}} = 1$$

Weight of $CaCO_3$ = no. of moles x molecular wt. = $1 \times 100 = 100$ g.

Q.2 What will be the number of oxygen atoms in 1 mole of O_2?

Sol. 1 mole of O_2 contains 6.023×10^{23} molecules of oxygen, and since the oxygen molecule is diatomic, i.e., 1 molecule contains 2 atoms, the no. of oxygen atoms in 1 mole of O_2 is equal to $2 \times 6.023 \times 10^{23}$.

Q.3 A piece of Cu weighs 0.635 g. How many atoms of Cu does it contain?

Sol. No of moles of Cu = wt. of Cu

at wt. of Cu $= \frac{0.635}{63.5} = 0.01$

No. of atoms of Cu= no. of moles of atoms × Av. Const.

$0.01 \times 6.023 \times 10^{23} = 6.023 \times 10^{21}$

Q.4 Calculate the number of molecules in 11.2 litres of SO_2 gas at NTP.

Sol. No. of moles of $SO_2 = \frac{\text{Volume at NTP (Litres)}}{\text{Standard molar volume (litres)}} = \frac{11.2}{22.4} = 0.5$

No. of molecules of SO_2 = No. of moles $\times$Av. Const.

$= 0.5 \times 6.022 \times 10^{23} = 3.011 \times 10^{3}$

Q.5 From 200 mg of CO_2, 10^{21} molecules are removed. How many moles of CO_2 are left?

Sol. Total no. of moles of $CO_2 = \frac{\text{wt. in g}}{\text{Mol. wt.}} = \frac{0.2}{44} = 0.004545$

No. of moles removed $= \frac{10^{21}}{6.022 \times 10^{23}} = 0.00166.$

No. of moles of CO_2 left = 0.004545–0.00166 = 0.00288

Q.6 Calculate the volume of 20 g of hydrogen gas at NTP.

Sol. Moles of hydrogen gas $\frac{20}{2} = 10$

Volume of the gas at NTP = No. of moles $\times 22.4$

$= 10 \times 22.4 = 224$ litres

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Q.7 Calculate the number of moles and the number of atoms of H, S and O in 5 moles of H_2SO_4.

Sol. 1 molecule of H_2SO_4 contains 2 atoms of H or 1 mole of H_2SO_4 contains 2 moles of H or 5 moles of H_2SO_4 contains 10 moles of H.

1 moles of H_2SO_4 contains 1 mole of S or 5 moles of H_2SO_4 contains 5 moles of S and again,

1 mole of H_2SO_4 contains 4 moles of O or 5 moles of H_2SO_4 contains 20 moles of O.

No. of atoms of H = $10 \times 6.022 \times 10^{23}$

No. of atoms of S = $5 \times 6.022 \times 10^{23}$

No. of atoms of O = $20 \times 6.022 \times 10^{23}$

Q.8 Calculate the volume occupied by 1 mole of He, H and O atoms at NTP.

Sol. As He is monoatomic, 1 mole of it will occupy 22.4 litres at NTP.

But hydrogen and oxygen being diatomic gases, 1 mole of their respective atoms will occupy a volume of 11.2 litres at NTP.

Salient features of Avogadro's hypothesis

Avogadro's law has played a very important role in the development of physical chemistry. The salient features of Avogadro's hypothesis are summarised below:

(i) It has removed the anomaly between Dalton's atomic theory and Gay Lussac's law of volume by making a clear distinction in between atoms and molecules.

(ii) It reveals that common elementary gases like hydrogen, nitrogen, oxygen, chlorine etc. are diatomic.

(iii) It provides a method to determine the atomic weights of gaseous elements,

(iv) It provides a relationship between vapour density and molecular weight of substances.

The vapour density of a gas is defined as the ratio of the weights of a certain volume of the gas to the weights of the same volume of hydrogen at same temperature and pressure. Thus,

$$\text{Vapor density} = \frac{\text{Volume of definite amount of gas}}{\text{Volume of same amount of hydrogen}}$$

$$\text{Vapor density} = \frac{\text{Weight of n molecules of gas}}{\text{Weight of n molecules of hydrogen}}$$

$$\text{Vapor density} = \frac{\text{Weight of one molecule of gas}}{\text{Weight of one molecule of hydrogen}}$$

$$\text{Vapor density} = \frac{\text{Weight of one molecule of gas}}{\text{Weight of one atom of hydrogen} \times 2}$$

$$\text{Molecular weight of a gas} = \frac{\text{Weight of one molecule of gas}}{\text{Weight of one atom of gas}}$$

$$\text{Vapor density} = \frac{\text{Molecular weight}}{2} \text{ or Molecular weight} = 2 \times \text{Vapor Density.}$$

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3.8 Chemical equation

A chemical equation is a short and quick way of representing a chemical reaction with the help of symbols and formulae of the reactants and products involved.

It tells us what substances entered into a given reaction (reactants) and what products were formed as a result of the reaction.

An arrow pointing towards the right as shown (i.e., towards the products) is placed between the reactants and the products.

Reactants → Products

This arrow shows that the substances written on the left (i.e., reactants) react together and produce the substances written on the right (i.e., products) of the arrow.

For example, When an aqueous solution of sodium hydroxide is added to that of copper sulphate, copper hydroxide is precipitated and sodium sulphate is formed in solution. The equation may be written like this:

$$\underset{\text{Reactants}}{2NaOH + CuSO_4} \rightarrow \underset{\text{Products}}{Cu(OH)_2 + Na_2SO_4}$$

Chemical reactions may involve any of these combinations:

(a) One reactant and two or more products
(b) Two reactants and one product
(c) Two reactants and two products
(d) Two reactants and three or more products.

(a) One reactant and two or more products:

(i) $CaCO_3 \longrightarrow CaO + CO_2$
(ii) $2Pb(NO_3)_2 \longrightarrow 2PbO + 4NO_2 + O_2$

(b) Two reactants and one product:

(i) $NH_3 + HCl \longrightarrow NH_4Cl$
(ii) $N_2 + 3H_2 \longrightarrow 2NH_3$

(c) Two reactants and two products:

(i) $AgNO_3 + NaCl \longrightarrow AgCl + NaNO_3$
(ii) $Na_2SO_4 + BaCl_2 \longrightarrow BaSO_4 + 2NaCl$

(d) Two reactants and three or more products:

(i) $Cu + 2H_2SO_4 \longrightarrow CuSO_4 + 2H_2O + SO_2$
(ii) $2KMnO_4 + 16\,HCl \longrightarrow 2KCl + 2MnCl_2 + 8H_2O + 5Cl_2$

Skeletal equation

This is an equation which represents a chemical change but is unbalanced. In other words, the total number of atoms of each element of the two sides is not equal.

For example, $KNO_3 \longrightarrow KNO_2 + O_2$

In the above example of the chemical equation, the number of oxygen atoms in the reactant (KNO_3) on the left is not equal to the number of atoms in the products formed ($KNO_2 + O_2$) on the right.

Balanced equation

In this equation the total number of atoms of each element in the reactants, on the left of the equation, is the same as the number of atoms in the products formed which is on the right of the equation. For example, $CaCO_3 \longrightarrow CaO + CO_2$

In this equation, the number of atoms of Ca, C and O on both sides is the same i.e., the equation is balanced.

$KNO_3 \longrightarrow KNO_2 + O_2$

In this equation, since the number of atoms of oxygen on both sides is not the same, so the equation is not balanced. The balanced form of the equation is:

$2KNO_3 \longrightarrow 2KNO_2 + O_2$

Why should an equation be balanced?

An equation must be balanced in order to follow the **Law of conservation of matter**. This law states that matter is neither created nor destroyed in the course of a chemical reaction. An unbalanced equation would imply that atoms have been created or destroyed.

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Information conveyed by a balanced chemical equation

Look at the following equation: $2NaOH + H_2SO_4 \longrightarrow Na_2SO_4 + 2H_2O$

This is a balanced chemical equation that gives us both **qualitative** and **quantitative** information.

(a) Qualitative information

A balanced qualitative chemical equation tells us the names, symbols and formulae of the reactant molecules taking part in the reaction and those of the product molecules formed in the reaction. In other words, a balanced qualitative chemical reaction tells us about the actual result of the chemical change (as in the above equation).

(b) Quantitative information

A balanced quantitative chemical equation tells us:

(a) Which substances take part in the reaction (i.e., the reactants) and which substances are formed (i.e., the products) as a result of it.

(b) The number of molecules of each substance taking part. Here. the two molecules of sodium hydroxide and one molecule of sulphuric acid react to give one molecule of sodium sulphate and two molecules of water.

(c) About the chemical composition of the respective molecules (about one molecule of sodium hydroxide contains one atom of sodium, one atom of oxygen and one atom of hydrogen).

(d) The mass of each substance involved in the reaction. In the above equation, sodium hydroxide reacts with sulphuric acid to produce sodium sulphate and water.

$$2NaOH + H_2SO_4 \longrightarrow Na_2SO_4 + 2H_2O$$

How can a chemical equation be made more informative?

(a) By writing **s, l, g** or **aq** to indicate the physical states of the reactants and the products involved in the chemical reaction. Solid, liquid, gaseous and aqueous solution states of the reactants and the products are indicated by writing **s** for solid, **l** for liquid, **g** for gaseous and **aq** for aqueous at the end of the symbol or formula of the reactants and the products.

For example, when solid zinc reacts with an aqueous solution of diluted sulphuric acid, zinc sulphate is obtained in solution and hydrogen gas is liberated. This reaction may be represented as:

$$Zn(s) + H_2SO_4\,(aq) \longrightarrow ZnSO_4\,(aq) + H_2\,(g)$$

The formation of a gaseous substance in a chemical reaction is indicated by putting an arrow pointing upwards with its formula (↑). Thus, the above reaction should be written as:

$$Zn(s) + H_2SO_4\,(aq) \longrightarrow ZnSO_4 + H_2\,(g)\uparrow$$

The formation of a precipitate (insoluble product) in a chemical reaction is indicated by putting an arrow facing downwards with its formula (↓).

For example, $Ca(OH)_2\,(aq) + CO_2\,(g) \xrightarrow{\text{Lime water}} H_2O(l) + CaCO_3\downarrow$

(b) By writing the heat term "+ heat" in the equation in which heat is either absorbed **(endothermic reaction)** or evolved **(exothermic reaction).** Chemical equations in which "+ heat" is included are known as **'thermochemical equations'**.

The term "+ heat" is written on the reactant side in the chemical equation in which heat energy is absorbed. For example, when carbon and sulphur are heated, carbon disulphide (CS_2) is formed. In this reaction, the heat is absorbed **(endothermic reaction).** The above reaction is thus represented by the chemical equation:

$$C\,(s) + 2S\,(g) + \text{heat } (= 92\text{ kJ}) \longrightarrow CS_2\,(g)$$

The term "+ heat" is written on the product side in the chemical equation in which heat energy is evolved **(exothermic reaction).** For example, when carbon is burnt in air or O_2, CO_2 is formed and heat is evolved. Thus, the above reaction is represented by the chemical equation.

$$C\,(s) + O_2\,(g) \longrightarrow CO_2\,(g) + \text{heat } (= 393\text{ kJ})$$

(c) By indicating the conditions under which the reaction takes place. If heat is required for a reaction to take place, delta (Δ) is written above the arrow (→) and if the reaction takes place in the presence of a catalyst, the symbol or formula of the catalyst is written above or below the arrow.

For example, when $KClO_3$ is heated in presence of MnO_2 (which is used as a catalyst), KCl and O_2 are obtained.

$$2KClO_3\,(s) \xrightarrow[MnO_2]{\Delta} 2KCl(s) + 3O_2\,(g)$$

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EXERCISE

Multiple choice questions

1. Two containers P and Q of equal volume (1 litre each) contain 6g of O_2 and SO_2 respectively at 300 K and 1 atmosphere. Then

(1) no. of molecules in P is less than that in Q.

(2) no. of the molecules in Q is less than that in P.

(3) no. of molecules in P and Q are same.

(4) either (1) or (2)

2. If nitrogen and ethylene are in the ratio 1: 2 by weight then their molecules are in the ratio of

(1) 1: 2
(2) 2: 1
(3) 1: 4
(4) 4: 1

3. The abbreviation used for the lengthy name of an element is termed as

(1) formulae
(2) symbol
(3) equation
(4) none of these

4. The number of atoms present in a molecule of a substance is called

(1) molecularity
(2) atomicity
(3) valency
(4) reactivity

5. Co stands for _____ while CO stands for ______ .

(1) the atoms of the element cobalt; carbon monoxide

(2) the atoms of the element; carbon monoxide

(3) the atom of the element cobalt; the molecules of the compound carbon monoxide

(4) the molecules and atoms of element carbon

6. Which of the following is the formula of the compound nickel bisulphate?

(1) $NiHSO_4$
(2) Ni_2HSO_4
(3) Ni_2SO_4
(4) $Ni(HSO_4)_2$

7. What is the chemical name of the substance whose formulae is $Na(NH_4)HPO_4$?

(1) Sodium hydrogen phosphate
(2) Ammonium hydrogen phosphate
(3) Sodium ammonium hydrogen phosphate
(4) None of these

8. Which of the following is the formula of the compound stannic phosphate?

(1) $Sn_3(PO_4)_4$
(2) $Sn_2(PO_3)_2$
(3) $Sn_3(PO_3)_2$
(4) $Sn_2(PO_3)_4$

9. Which of the following is the formula of the compound magnesium phosphite?

(1) Mg_2PO_3
(2) $Mg_2(PO_3)_4$
(3) $MgHPO_3$
(4) $Mg_3(PO_3)_2$

10. Which of the following is the chemical name of $Ba(ClO_3)_2$?

(1) Barium chloride
(2) Barium chlorate
(3) Barium chlorite
(4) Barium hypochlorite

11. Which of the following is the formula of barium peroxide?

(1) Ba_2O
(2) Ba_2O_2
(3) BaO_2
(4) BaO_3

12. The molecular formula of ethanoic acid is $C_2H_4O_2$. Its empirical formula is

(1) $C_4H_8O_2$
(2) CH_2O
(3) CHO
(4) CHO_2

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13. If we consider that $\frac{1}{6}$, in place of $\frac{1}{12}$, mass of carbon atom is taken to be relative atomic mass unit, the mass of one mole of a substance will

(1) increase two fold.
(2) decrease twice.
(3) be a function of molecular mass of the substance.
(4) remain unchanged.

14. Chemical analysis of a carbon compound gave the following percentage composition by weight of the elements present in it. Carbon = 10.06%, hydrogen = 0.84%, chlorine = 89.10%. Calculate the empirical formula of the compound.

(1) $C_2H_2Cl_2$
(2) $CHCl_2$
(3) $CHCl_3$
(4) $C_4H_4Cl_4$

15. 0.202 g of a carbon compound, on combustion, gave 0.361 g of carbon dioxide and 0.47 g of water. Calculate the percentage composition of carbon.

(1) 48.73%
(2) 8.07%
(3) 43.17%
(4) 42.17%

16. The percentage of nitrogen in ammonia is given by the expression

(1) $\frac{14\times100}{17}$
(2) $\frac{3}{17}\times100$
(3) $\frac{14\times100}{34}$
(4) $\frac{3}{34}\times100$

17. The empirical formula of a compound is CH_2O. Its molecular weight is 90. Calculate the molecular formula of the compound. (Atomic weight C = 12, H = 1, O = 16)

(1) $C_3H_7O_3$
(2) $C_4H_6O_3$
(3) $C_3H_6O_3$
(4) $C_4H_2O_3$

18. A compound contains 38.8% C, 16% H and 45.2% N. The empirical formula of the compound is

(1) CH_3NH_2
(2) CH_2CN
(3) C_2H_5CN
(4) $CH_2(NH_2)_2$

19. A hydrocarbon contains 90% of carbon and 10% hydrogen. The empirical formula of the compound is

(1) C_2H_5
(2) C_3H_2
(3) C_3H_4
(4) CH_3

20. 64 g of an organic compound contains 24 g of carbon, 8 g of hydrogen and the rest oxygen. The empirical formula of the compound is

(1) CH_2O
(2) C_2H_4O
(3) CH_4O
(4) $C_2H_4O_2$

21. Percentage of carbon is highest in

(1) CH_4
(2) C_3H_4
(3) C_6H_6
(4) C_3H_8

22. Find out the percentage of carbon in aluminium carbonate $Al_2(CO_3)_3$ [Al = 27, C = 12, O = 16].

(1) 15.38%
(2) 14.38%
(3) 27.2%
(4) 12.16%

23. How many litres of steam will be formed from 2L of H_2 and 1 L of O_2, if all volumes are measured at the same temperature and pressure?

(1) 1 L
(2) 2 L
(3) 3 L
(4) 4 L

24. The number of molecules in 35.5 g of chlorine gas is

(1) 3.011×10^{23}
(2) 6.022×10^{23}
(3) 9.033×10^{23}
(4) 1.2044×10^{24}

25. Which of the following has the largest number of atoms? [Cu = 63.5 u]

(1) 0.5 g atoms of Cu
(2) 0.635 g of Cu
(3) 0.25 moles of Cu atoms
(4) 1 g of Cu

26. Total number of electrons present in 18 ml of water (density of water is 1 g mL^{-1}).

(1) 6.02×10^{23}
(2) 6.02×10^{22}
(3) 6.02×10^{24}
(4) 6.02×10^{25}

27. The number of moles of sodium oxide in 620 g of it is

(1) 1 mole
(2) 10 moles
(3) 18 moles
(4) 100 moles

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28. Volume of gas at STP is 1.12×10^{-7} cc. Calculate the no. of molecules in it

(1) 3.01×10^{20} (2) 3.01×10^{12} (3) 3.01×10^{23} (4) 3.01×10^{24}

29. How many moles of electron weigh one kilogram?

(1) 6.023×10^{23} (2) $\frac{1}{9.108} \times 10^{31}$ (3) $\frac{6.023}{9.108} \times 10^{54}$ (4) $\frac{1}{9.108 \times 6.023} \times 10^{8}$

30. Arrange the following in order of increasing mass (atomic mass: O = 16, Cu = 63, N = 14)

I. One atom of oxygen II. One atom of nitrogen

III. 1×10^{-10} mole of oxygen IV. 1×10^{-10} mole of copper

(1) II < I < III < IV (2) I < II < III < IV

(3) III < II < IV < I (4) IV < II < III < I

31. 25.3 g of sodium carbonate, Na_2CO_3 is dissolved in enough water to make 250 mL of solution. If sodium carbonate dissociates completely, molar concentration of sodium ions, Na^+ and carbonate ions, CO_3^{2-} are respectively (Molar mass of Na_2CO_3 = 106 g mol^{-1})

(1) 0.477 M and 0.477 M (2) 0.955 M and 1.910 M

(3) 1.910 M and 0.955 M (4) 1.90 M and 1.910 M

32. The crystalline salt $Na_2SO_4 . x\ H_2O$ on heating loses 55.9% of its weight. The formula of the crystalline salt is

(1) $Na_2SO_4 . 5\ H_2O$ (2) $Na_2SO_4 . 7\ H_2O$

(3) $Na_2SO_4 . 2\ H_2O$ (4) $Na_2SO_4 . 10\ H_2O$

33. What is the weight of 3 gram atoms of sulphur?

(1) 96 gm (2) 99 gm (3) 100 gm (4) 3 gm

34. How many moles of oxygen atoms are present in one mole of acetic acid?

(1) 1 mole (2) 3 moles (3) 2 moles (4) 6 moles

35. Calculate the weight of 0.1 mole of sodium carbonate.

(1) 1.06 g (2) 11.06 g (3) 10.6 g (4) 1.106 g

36. How much of lime can be obtained by burning 400 g of lime stone?

(1) 224 g (2) 220 g (3) 400 g (4) 320 g

37. 3.6 grams of a sample, on combustion, gave 3.3 grams of CO_2. The percentage of carbon in the sample is

(1) 10% (2) 20% (3) 25% (4) 40%

38. The number of moles present in 20 grams of $CaCO_3$ is

(1) 0.1 (2) 0.2 (3) 0.3 (4) 0.25

39. Calculate the number of iron atoms in a piece of iron weighing 2.8 g. (Atomic mass of iron = 56)

(1) 30.11×10^{23} atoms (2) 3.11×10^{23} atoms

(3) 3.011×10^{22} atoms (4) 301.1×10^{23} atoms

40. Sulphur molecule exists as S_8. How many moles of sulphur are present in 25.6 grams of sulphur?

(1) 0.01 (2) 0.1 (3) 0.02 (4) 0.2

41. How many grams of sodium oxalate are to be weighed to get 0.1 mole of sodium oxalate ($Na_2C_2O_4$)?

(1) 13.4 g (2) 1.34 g (3) 134 g (4) 0.134 g

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42. Calculate the number of gram atoms and gram molecules present in 12.4 g of phosphorus.

(1) 0.1 gram atoms, 0.1 gram molecules
(2) 0.2 gram atoms, 0.1 gram molecules
(3) 0.1 gram atoms, 0.4 gram molecules
(4) 0.4 gram atoms, 0.1 gram molecules

43. Calculate the number of moles present in 4.9 g of H_2SO_4.

(1) 0.01
(2) 0.02
(3) 0.05
(4) 0.03

44. How many atoms and gram atoms are present in 20 g of calcium?

(1) 3.1×10^{23} and 0.05
(2) 3.01×10^{23} and 0.5
(3) 0.5 and 3.1×10^{23}
(4) 0.2 and 3.1×10^{23}

45. Calculate the weight of nitrogen present in 0.5 moles of NH_3.

(1) 8 g
(2) 9 g
(3) 1 g
(4) 7 g

46. Calculate the mass of one molecule of H_2O.

(1) 29.88×10^{23} g
(2) 2.988×10^{23} g
(3) 0.2988×10^{23} g
(4) 2.988×10^{-23} g

47. Calculate the weight of carbon which contains the same number of atoms as those present in 15 grams of oxygen.

(1) 13 g
(2) 11.25 g
(3) 10 g
(4) 14 g

48. What weight of sodium contains the same number of atoms as those in 8 grams of oxygen?

(1) 10.5 g
(2) 13.5 g
(3) 11.5 g
(4) 14.2 g

49. How many atoms are present in 0.25 mole of SO_3?

(1) 1.505×10^{23} atoms of sulphur and 4.515×10^{23} atoms of oxygen
(2) 4.15×10^{23} atoms of oxygen and 2.505×10^{22} atoms of sulphur
(3) 2.505×10^{23} atoms of sulphur and 3.505×10^{22} atoms of oxygen
(4) 3.05×10^{22} atoms of sulphur and 2.515×10^{23} atoms of oxygen

50. How many molecules are present in 0.50 moles of KCl?

(1) 4.01×10^{23}
(2) 3.01×10^{22}
(3) 3.01
(4) 3.01×10^{23}

51. Calculate the number of gram atoms present in 103 g of barium. [Ba = 137.4]

(1) 103
(2) 0.4
(3) 0.75
(4) 0.3

52. Calculate the weight in gram present in 0.7 moles of sodium.

(1) 16.1 gram
(2) 16.2 gram
(3) 16.3 gram
(4) 0.161 gram

53. Calculate the weight of 0.8 moles of magnesium.

(1) 20 g
(2) 19.2 g
(3) 30 g
(4) 31.2 g

54. Calculate the weight of 10^{24} molecules of nitric oxide [NO].

(1) 49.8 g
(2) 50 g
(3) 50.1 g
(4) 39.4 g

55. Calculate the total number of atoms present in 0.8 moles of H_2S.

(1) 14.4
(2) 14.4×10^{22}
(3) 14.452×10^{23}
(4) 14.4×10^{-23}

56. Calculate the number of oxygen atoms present in 0.75 moles of SO_2.

(1) 9.03×10^{23}
(2) 9.03×10^{22}
(3) 9.03
(4) 9.03×10^{-23}

57. If a mole of oxygen contains 1.00×10^{24} particles, calculate the mass of a single oxygen molecule.

(1) 32×10^{-24} g
(2) 32×10^{-23} g
(3) 32 g
(4) 32×10^{24} g

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58. Calculate the number of moles and molecules present in 4 grams of methane.

(1) 0.25 moles, 1.505 × 10^{23} molecules

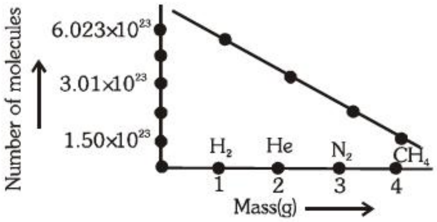
(2) 0.3 moles, 1.505 molecules

(3) 0.25 moles, 1.505 × 10^{-23} molecules

(4) 0.3 moles, 2.0 molecules

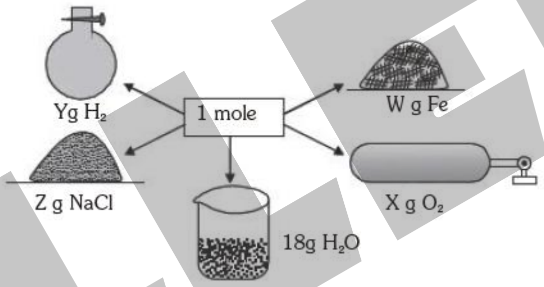
59. Given is the graphical representation of number of molecules of different gases.

Which is/are placed at correct position?



(1) H_2, He (2) N_2, CH_4 (3) He, CH_4 (4) H_2, CH_4

60. Observe the given figure carefully and select the correct option.



(1) W = 55.8, X = 16, Y = 2.016, Z = 58.5

(2) W = 55.8, X = 32, Y = 1.008, Z = 58.5

(3) W = 55.8, X = 16, Y = 1.008, Z = 58.5

(4) W = 55.8, X = 32, Y = 2.016, Z = 58.5

ANSWERS

Que.	**1**	**2**	**3**	**4**	**5**	**6**	**7**	**8**	**9**	**10**	**11**	**12**	**13**	**14**	**15**	**16**	**17**	**18**	**19**	**20**
Ans.	2	1	2	2	3	4	3	1	4	2	3	2	4	3	1	1	3	1	3	3
Que.	**21**	**22**	**23**	**24**	**25**	**26**	**27**	**28**	**29**	**30**	**31**	**32**	**33**	**34**	**35**	**36**	**37**	**38**	**39**	**40**
Ans.	3	1	2	1	1	3	2	2	4	1	3	4	1	3	3	1	3	2	3	2
Que.	**41**	**42**	**43**	**44**	**45**	**46**	**47**	**48**	**49**	**50**	**51**	**52**	**53**	**54**	**55**	**56**	**57**	**58**	**59**	**60**
Ans.	1	4	3	2	4	4	2	3	1	4	3	1	2	1	3	1	1	1	3	4

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